

Chapter 4 – Bonding and Structure

Ionic bonding

- Ionic structures
- Properties of ionic compounds
- Formulae of ionic substances

Covalent bonding

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- Covalent structures
- Properties of covalent compounds
- Allotropy
- Allotropes of carbon

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Additional questions

In Chapter 3 we have seen that the noble gases are stable or unreactive because they have full electron energy levels (Figure 4.1). It would seem that when elements react to form compounds they do so to achieve full electron energy levels. This idea forms the basis of the electronic theory of chemical bonding.



Figure 4.1 The unreactive noble gas neon is used in lasers.

Ionic bonding

Ionic bonds are usually found in compounds that contain metals combined with non-metals. When this type of bond is formed, electrons are transferred from the metal atoms to the non-metal atoms during the chemical reaction. In doing this, the atoms become more stable by getting full outer energy levels. For example, consider what happens when sodium and chlorine combine to make sodium chloride (Figure 4.2).

sodium + chlorine - sodium chloride



Figure 4.2 The properties of salt are very different from those of the sodium and chlorine it was made from. To get your salt you would not eat sodium or inhale chlorine.

Sodium has just one electron in its outer energy level (^{11}Na 2,8,1). Chlorine has seven electrons in its outer energy level (^{17}Cl 2,8,7). When these two

elements react, the outer electron of each sodium atom is transferred to the outer energy level of a chlorine atom (Figure 4.3). In this way both the atoms obtain full outer energy levels and become 'like' the nearest noble gas. The sodium atom has become a sodium ion with an electron configuration like neon, while the chlorine atom has become a chloride ion with an electron configuration like argon. Only the outer electrons are important in bonding, so we can simplify the diagrams by missing out the inner energy levels (Figure 4.4).

The charges on the sodium and chloride ions are equal but opposite. They balance each other and the resulting formula for sodium chloride is NaCl. These oppositely charged ions attract each other and are pulled, or **bonded**, to one another by strong electrostatic forces. This type of bonding is called **ionic bonding**. The alternative name, **electrovalent bonding**, is derived from the fact that there are electrical charges on the atoms involved in the bonding.

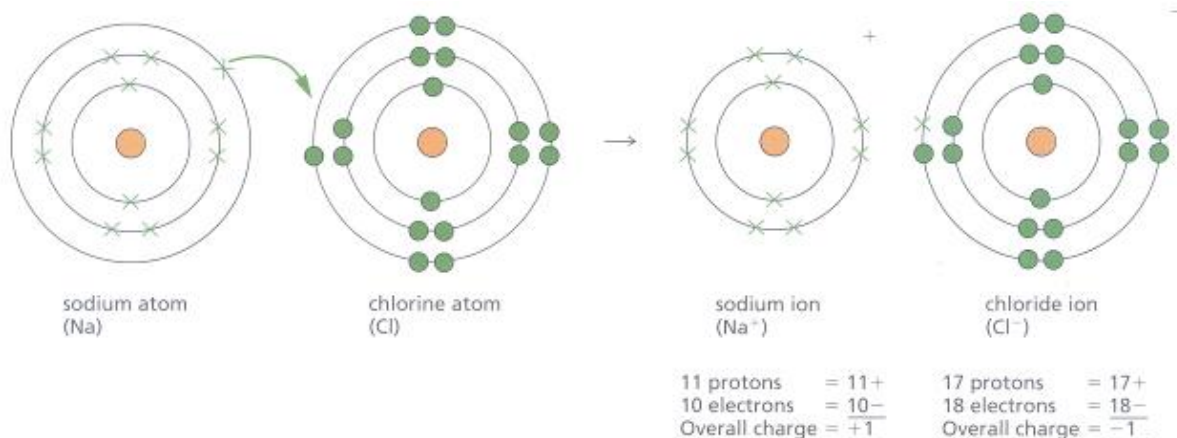


Figure 4.3 Ionic bonding in sodium chloride.

Figure 4.5 shows the electron transfers that take place between a magnesium ion and an oxide ion during the formation of magnesium oxide.

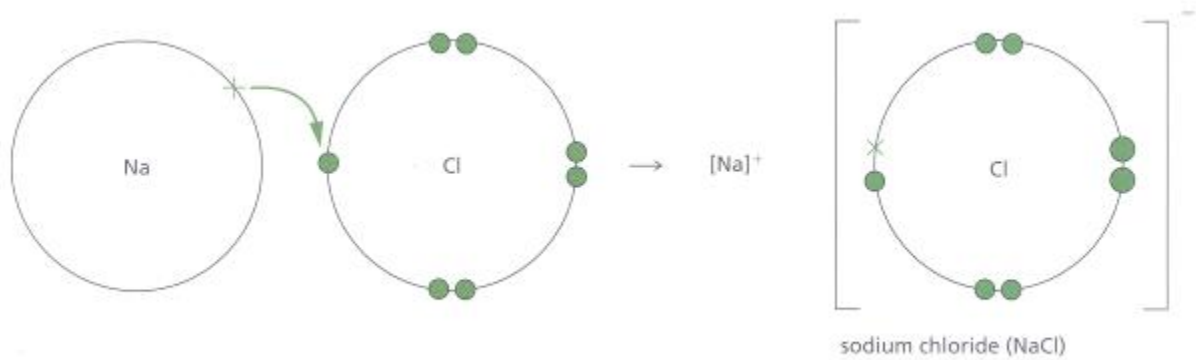


Figure 4.4 Simplified diagram of ionic bonding in sodium chloride.

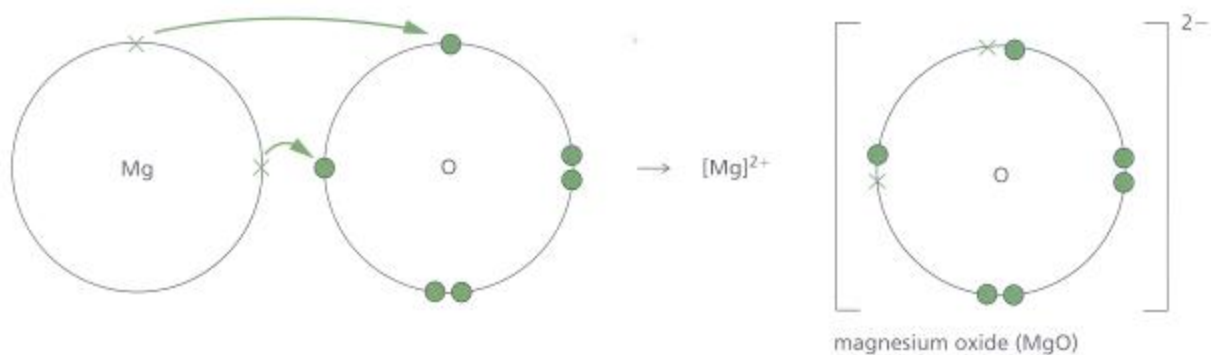


Figure 4.5 Simplified diagram of ionic bonding in magnesium oxide.

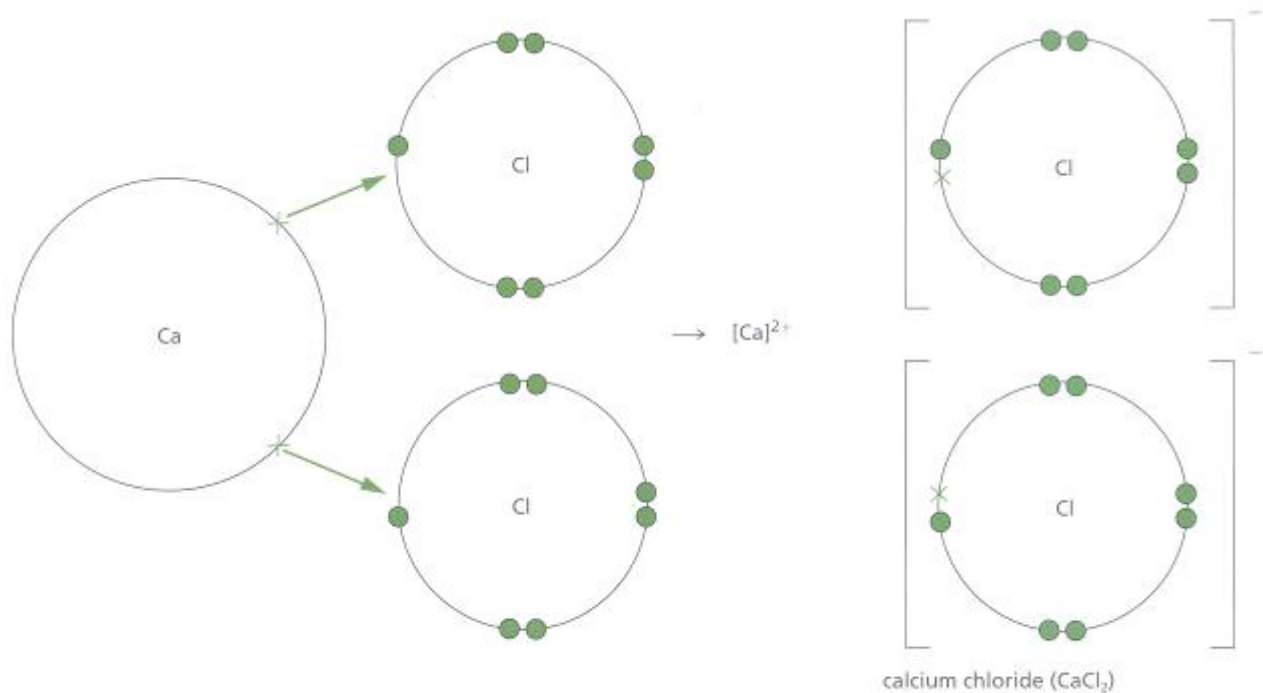


Figure 4.6 The transfer of electrons that takes place during the formation of calcium chloride.

Magnesium obtains a full outer energy level by losing two electrons. These are transferred to the oxygen atom. In magnesium oxide, the Mg^{2+} and O^{2-} are oppositely charged and are attracted to one another. The formula for magnesium oxide is MgO .

Figure 4.6 shows the electron transfers that take place during the formation of calcium chloride. When these two elements react, the calcium atom gives each of the two chlorine atoms one electron. In this case, a compound is formed containing two chloride ions (Cl^-) for each calcium ion (Ca^{2+}). The chemical formula is CaCl_2 .

Question

1. Draw diagrams to represent the bonding in each of the following ionic compounds:

- a. magnesium fluoride (MgF_2)
- b. potassium fluoride (KF)
- c. lithium chloride (LiCl)
- d. calcium oxide (CaO).

Ionic structures

Ionic structures are solids at room temperature and have high melting and boiling points. The ions are packed together in a regular arrangement called a **lattice**. Within the lattice, oppositely charged ions attract one another strongly. Scientists, using **X-ray** diffraction (Figure 4.7a), have obtained photographs that indicate the way in which the ions are arranged (Figure 4.7b). The electron density map of sodium chloride is shown in Figure 4.7c.

Figure 4.7d shows the structure of sodium chloride as determined by the X-ray diffraction technique. The study of crystals using X-ray diffraction was pioneered by Sir William Bragg and his son Sir Lawrence Bragg in 1912. X-rays are a form of electromagnetic radiation. They have a much shorter wavelength than light therefore it is possible to use them to investigate extremely small structures.

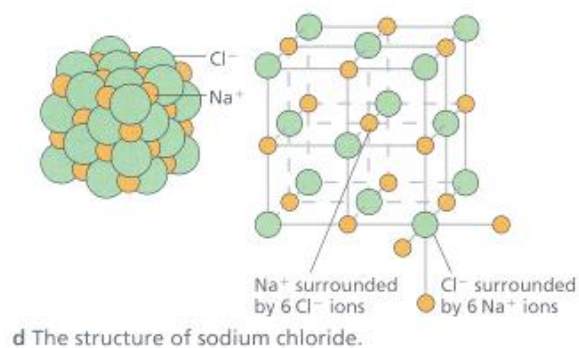
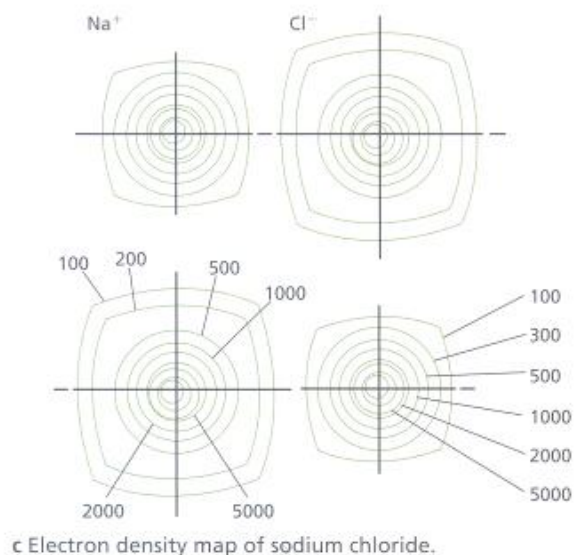
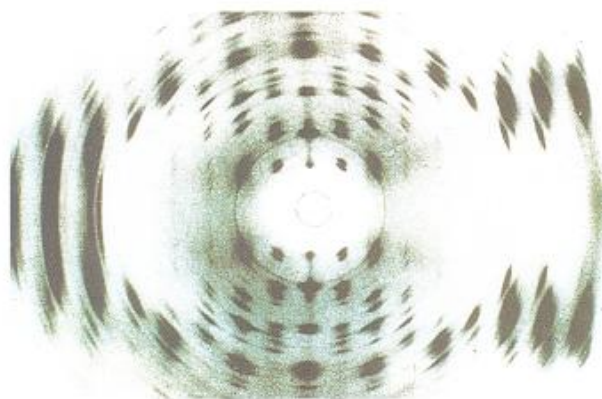
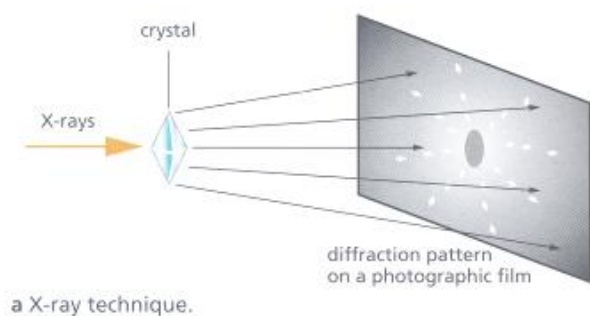


Figure 4.7

When X-rays are passed through a crystal of sodium chloride, for example, you get a pattern of spots called a diffraction pattern (Figure 4.7b). This pattern can be recorded on photographic film and used to work out how the ions or atoms are arranged in the crystal. Crystals give particular diffraction patterns depending on their structure, and this makes X-ray diffraction a particularly powerful technique in the investigation of crystal structures.

Figure 4.7d shows only a tiny part of a small crystal of sodium chloride. Many millions of sodium ions and chloride ions would be arranged in this way in a crystal of sodium chloride to make up the giant ionic structure. Each sodium ion in the lattice is surrounded by six chloride ions, and each chloride ion is surrounded by six sodium ions.

Not all ionic substances form the same structures. Caesium chloride (CsCl), for example, forms a different structure due to the larger size of the caesium ion compared with that of the sodium ion. This gives rise to the structure shown in Figure 4.8, which is called a body-centred cubic structure. Each caesium ion is surrounded by eight chloride ions and, in turn, each chloride ion is surrounded by eight caesium ions.

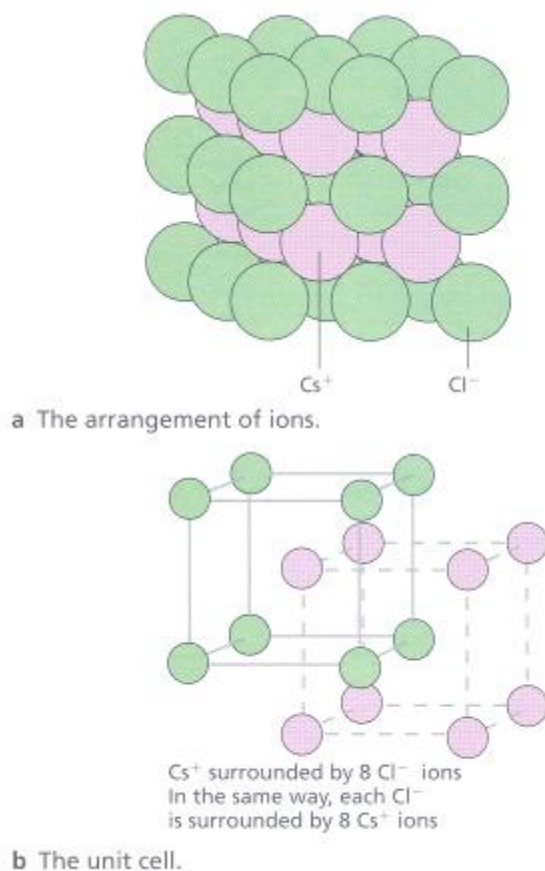


Figure 4.8 Structure of caesium chloride.

Properties of ionic compounds

Ionic compounds have the following properties.

- They are usually solids at room temperature, with high melting points. This is due to the strong electrostatic forces holding the crystal lattice together. A lot of energy is therefore needed to separate the ions and melt the substance.
- They are usually hard substances.
- They usually cannot conduct electricity when solid, because the ions are not free to move.
- They mainly dissolve in water. This is because water molecules are able to bond with both the positive and the negative ions, which breaks up the lattice and keeps the ions apart.

- They usually conduct electricity when in the molten state or in aqueous solution. The forces of attraction between the ions are weakened and the ions are free to move. This allows an electric current to be passed through the molten sodium chloride. For a further discussion of this process see Chapter 6.

Formulae of ionic substances

We saw that ionic compounds contain positive and negative ions, whose charges balance. For example, sodium chloride contains one Na^+ ion for every Cl^- ion, giving rise to the formula NaCl . This method can be used to write down formulae which show the ratio of the number of ions present in any ionic compound.

The formula of magnesium chloride is MgCl_2 . This formula is arrived at by each Mg^{2+} ion combining with two Cl^- ions, and once again the charges balance. The size of the charge on an ion is a measure of its **valency** or **combining power**. Na^+ has a valency of 1, but Mg^{2+} has a valency of 2. Na^+ can bond (combine) with only one Cl^- ion, whereas Mg^{2+} can bond with two Cl^- ions.

Some elements, such as copper and iron, possess two ions with different valencies. Copper can form the Cu^+ ion and the Cu^{2+} ion, and therefore it can form two different compounds with chlorine, CuCl (copper (I) chloride) and CuCl_2 (copper (II) chloride). Iron forms the Fe^{2+} and Fe^{3+} ions.

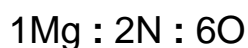
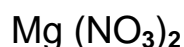
Table 4.1 shows the valencies of a series of ions you will normally meet in your study of chemistry.

Table 4.1 Valencies of some common substances.

	Valency		
	1	2	3
Metals	Lithium (Li^+)	Magnesium (Mg^{2+})	Aluminium (Al^{3+})
	Sodium (Na^+)	Calcium (Ca^{2+})	Iron (Fe^{3+})
	Potassium (K^+)	Copper (Cu^{2+})	
	Silver (Ag^+)	Zinc (Zn^{2+})	
	Copper (Cu^+)	Iron (Fe^{2+})	

		Lead (Pb^{2+})	
		Barium (Ba^{2+})	
Non-metals	Fluoride (F^-)	Oxide (O^{2-})	
	Chloride (Cl^-)	Sulphide (S^{2-})	
	Bromide (Br^-)		
	Hydrogen (H^+)		
Groups of atoms	Hydroxide (OH^-)	Carbonate (CO_3^{2-})	Phosphate (PO_4^{3-})
	Nitrate (NO_3^-)	Sulphate (SO_4^{2-})	
	Ammonium (NH_4^+)		
	Hydrogen carbonate (HCO_3^-)		

You will notice that Table 4.1 includes groups of atoms which have net charges. For example, the nitrate ion is a single unit composed of one nitrogen atom and three oxygen atoms and has one single negative charge. The formula, therefore, of magnesium nitrate would be $\text{Mg}(\text{NO}_3)_2$. You will notice that the NO_3 has been placed in brackets with a 2 outside the bracket. This indicates that there are two nitrate ions present for every magnesium ion. The ratio of the atoms present is therefore:



Questions

1. Using the information in Table 4.1, write the formulae for:

- copper (I) oxide
- zinc phosphate
- iron (III) chloride
- lead bromide.

2. Using the formulae in your answer to question 1, write down the ratio of atoms present for each of the compounds.

Covalent bonding

Another way in which atoms can gain the stability of the noble gas electron configuration is by sharing the electrons in their outer energy levels. This occurs between non-metal atoms, and the bond formed is called a **covalent bond**. The simplest example of this type of bonding can be seen by considering the hydrogen molecule, H_2 .

Each hydrogen atom in the molecule has one electron. In order to obtain a full outer energy level and gain the electron configuration of the noble gas helium each of the hydrogen atoms must have two electrons. To do this, the two hydrogen atoms allow their outer energy levels to overlap (Figure 4.9). A molecule of hydrogen is formed, with two hydrogen atoms sharing a pair of electrons (Figure 4.10). This shared pair of electrons is known as a single covalent bond and is represented by a single line as in hydrogen:

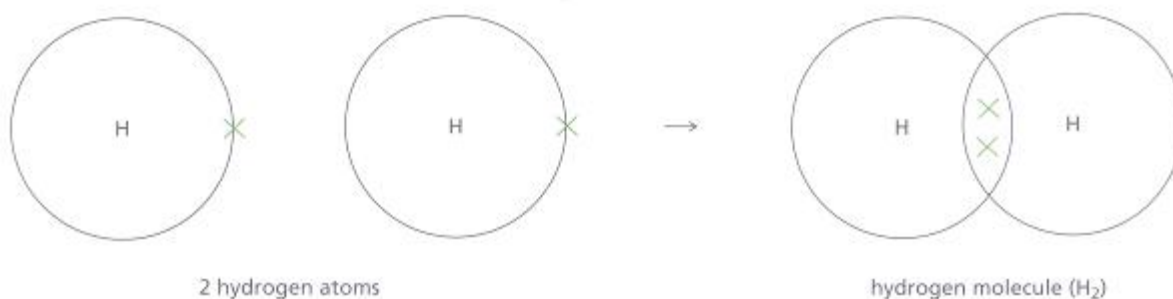
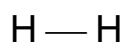
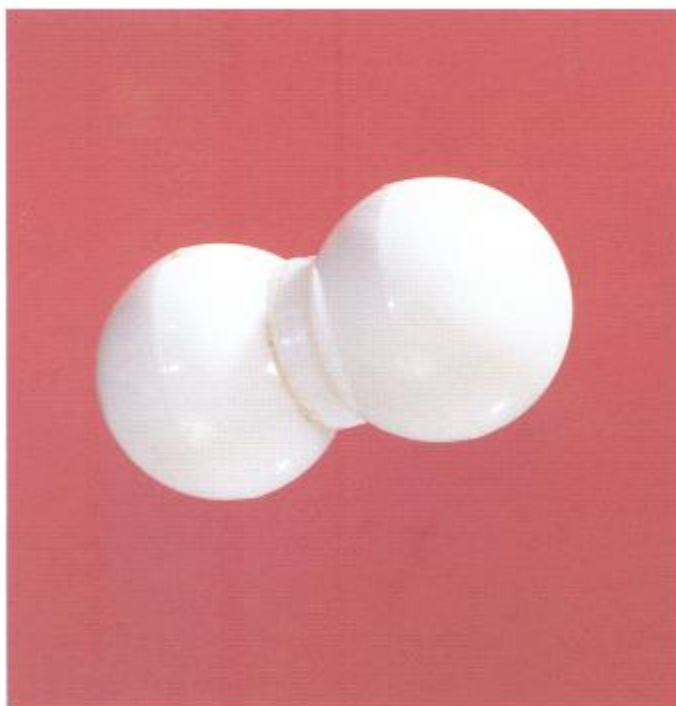
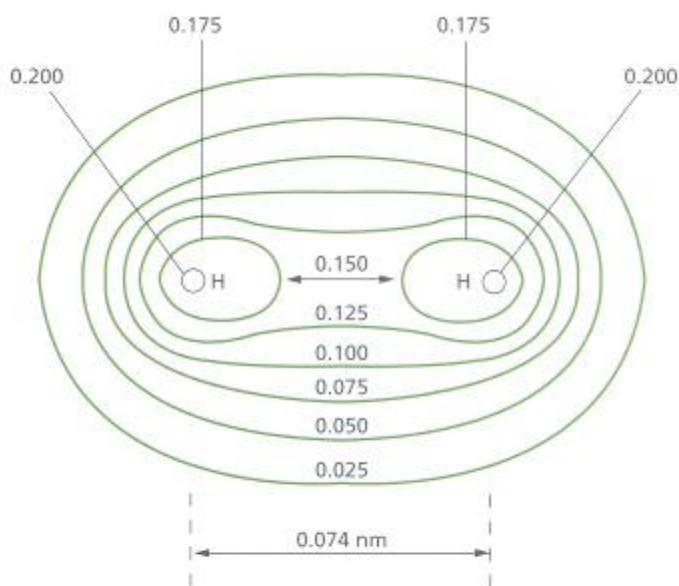


Figure 4.9 Electron sharing to form the single covalent bond in H_2 molecules.



a Model of hydrogen molecule.

Figure 4.10



electron densities in electrons per cubic atomic unit of length

b The electron density map of a hydrogen molecule.

Other covalent compounds

Methane (natural gas) is a gas whose molecules contain atoms of carbon and hydrogen. The electron structures are:



The carbon atom needs four more electrons to attain the electron configuration of the noble gas neon. Each hydrogen atom needs only one electron to form the electron configuration of helium. Figure 4.11 shows how the atoms gain these electron configurations by the sharing of electrons. You will note that only the outer electron energy levels are shown. Figure 4.12 shows the shape of the methane molecule.

Ammonia is a gas containing the elements nitrogen and hydrogen. It is used in large amounts to make fertilisers. The electron configurations of the two elements are:



The nitrogen atom needs three more electrons to obtain the noble gas structure of neon. Each hydrogen requires only one electron to form the noble gas structure of helium. The nitrogen and hydrogen atoms share electrons, forming three single covalent bonds (Figure 4.13). Unlike methane the shape of an ammonia molecule is pyramidal (Figure 4.14).

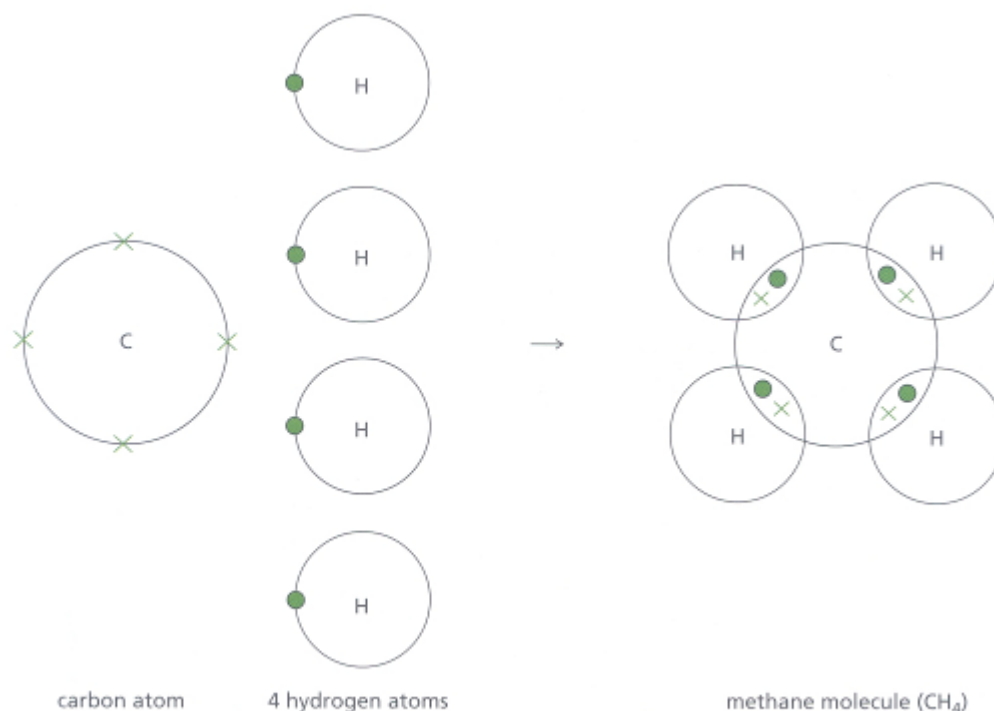
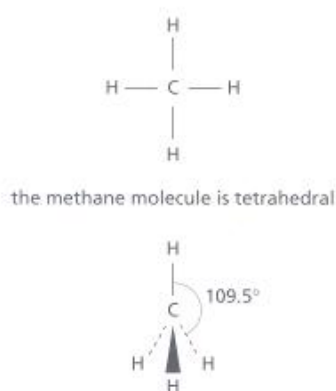


Figure 4.11 Formation of methane.



a Methane molecule.

Figure 4.12



b Model of the methane molecule. The tetrahedral shape is caused by the repulsion of the C–H bonding pairs of the electrons.

Water is a liquid containing the elements hydrogen and oxygen. The electron configurations of the two elements are:

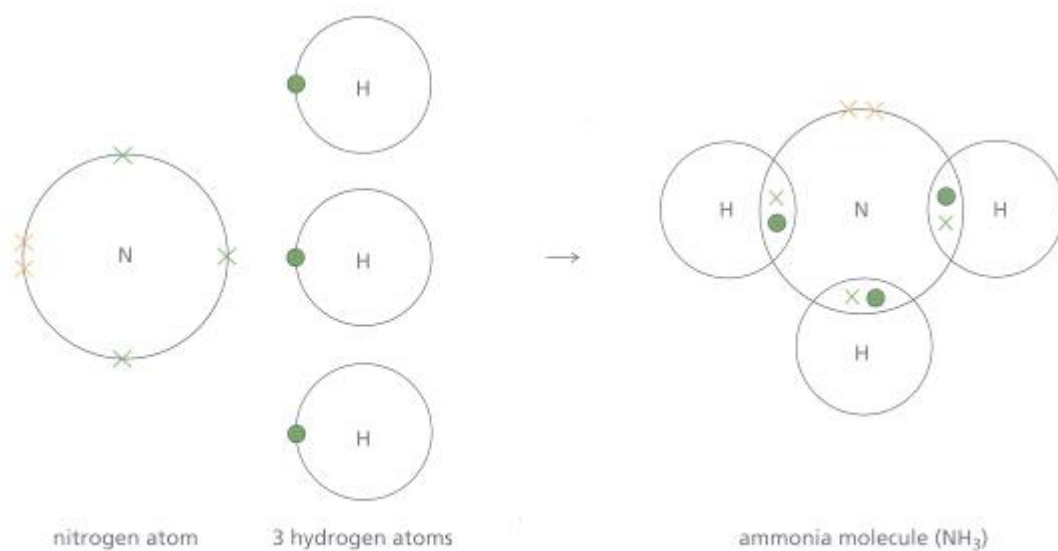
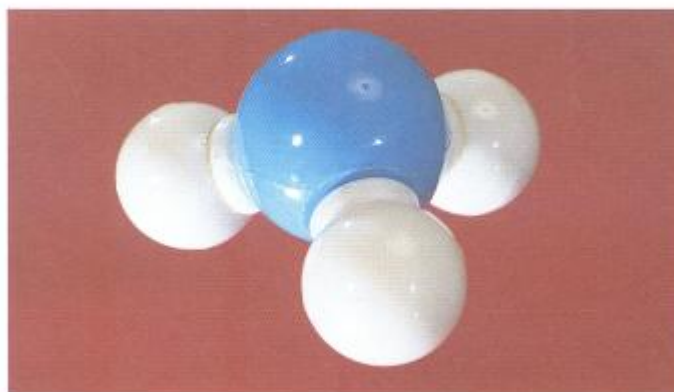


Figure 4.13 The bonding in ammonia.



a Ammonia molecule.



b Model of the ammonia molecule. The pyramidal shape is caused by the repulsion between the bonding pairs of electrons as well as the lone pair (or non-bonding pair) of electrons.

Figure 4.14

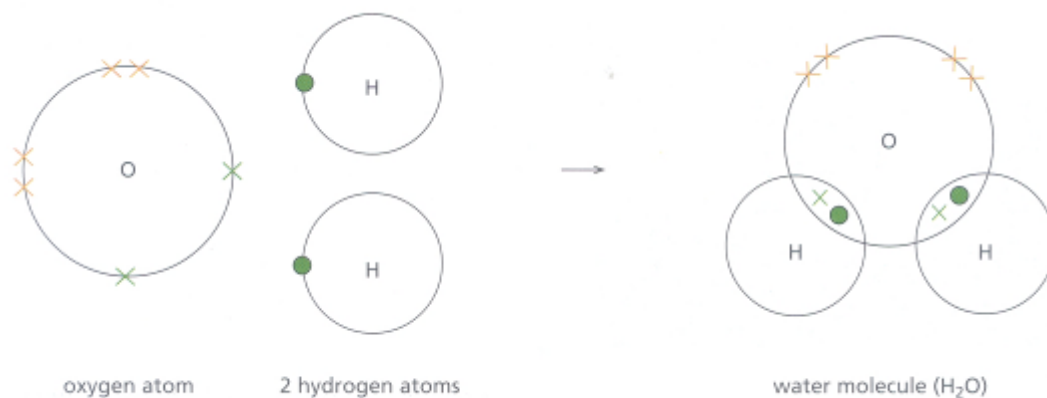
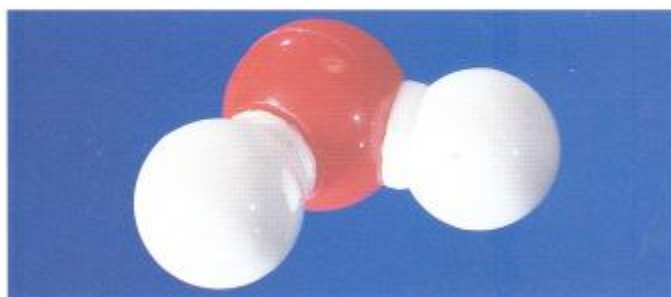


Figure 4.15 Formation of water.

The oxygen atom needs two electrons to gain the electron configuration of neon. Each hydrogen requires one more electron to gain the electron configuration of helium. Again, the oxygen and hydrogen atoms share electrons, forming a water molecule with two single covalent bonds as shown in Figure 4.15. A water molecule is V-shaped (Figure 4.16).



a Water molecule.



b Model of a water molecule. This is a V-shaped molecule.

Figure 4.16

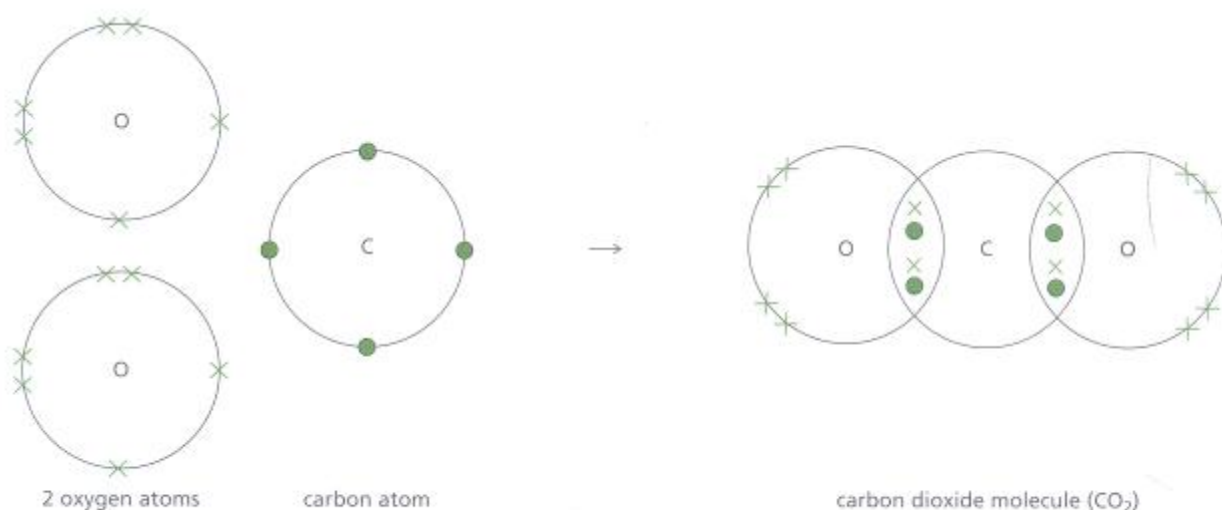


Figure 4.17 Formation of carbon dioxide.

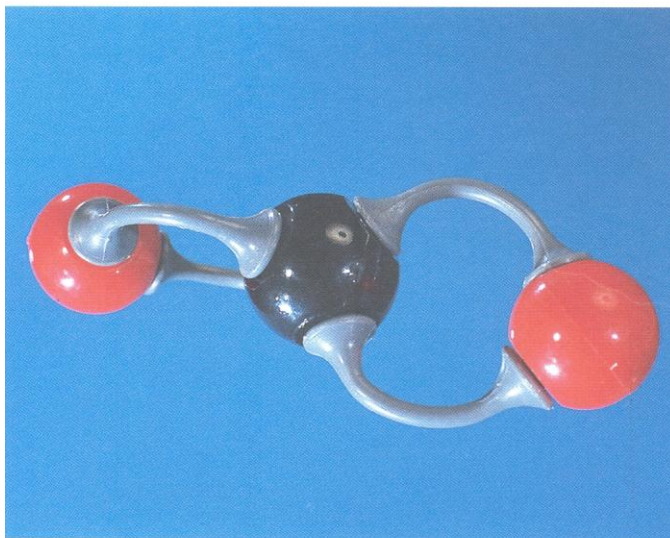
Carbon dioxide is a gas containing the elements carbon and oxygen. The electron configurations of the two elements are:



In this case each carbon atom needs to share four electrons to gain the electron configuration of neon. Each oxygen needs to share two electrons to gain the electron configuration of neon. This is achieved by forming two **double covalent bonds** in which two pairs of electrons are shared in each case, as shown in Figure 4.17. Carbon dioxide is a linear molecule (Figure 4.18).



a Carbon dioxide molecule. Note the double covalent bond is represented by a double line.



b Model of the linear carbon dioxide molecule.

Figure 4.18

Questions

1 Draw diagrams to represent the bonding in each of the following covalent compounds:

- a. tetrachloromethane (CCl_4)
- b. oxygen gas (O_2)
- c. hydrogen sulphide (H_2S)
- d. hydrogen chloride (HCl)
- e. ethene (C_2H_4)
- f. methanol (CH_3OH)
- g. nitrogen (N_2).

2. Explain why the water molecule in Figure 4.16 is V-shaped.

Covalent structures

Compounds containing covalent bonds have molecules whose structures can be classified as either **simple molecular** or **giant molecular**.

Simple molecular structures are simple, formed from only a few atoms. They have strong covalent bonds between the atoms within a molecule (**intramolecular bonds**) but have weak bonds between the molecules (intermolecular bonds) (Figure 4.19). One type of weak bond between molecules is known as the **van der Waals' bond** (or force), and these forces increase steadily with the increasing size of the molecule. Examples of simple molecules are iodine, methane, water and ethanol.

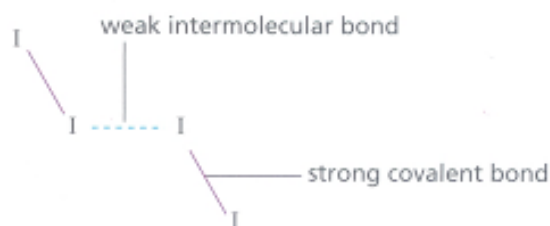
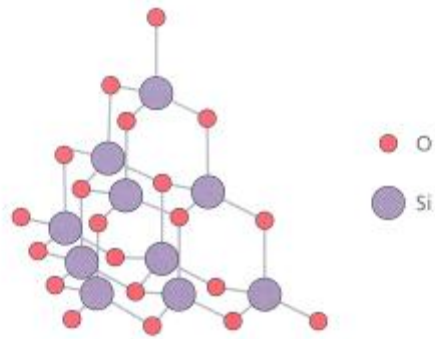


Figure 4.19 Strong covalent and weak intermolecular forces in iodine.

Giant molecular structures contain many hundreds of thousands of atoms joined by strong covalent bonds. Examples of substances showing this type of structure are diamond, graphite, silicon (IV) oxide (Figure 4.20) and plastics such as polythene (polymers, Chapter 14). Plastics are a tangled mass of very long molecules in which the atoms are joined together by strong covalent bonds to form long chains. Molten plastics can be made into fibres by being forced through hundreds of tiny holes in a 'spinneret' (Figure 4.21). This process aligns the long chains of atoms along the length of the fibre. For a further discussion of plastics, see Chapter 14.



a The silicon(IV) oxide structure in quartz.



b Quartz is a hard solid at room temperature. It has a melting point of 1610°C and a boiling point of 2230°C .

Figure 4.20

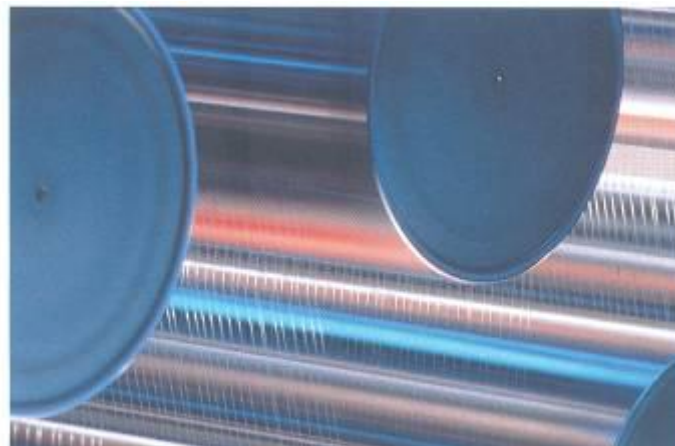
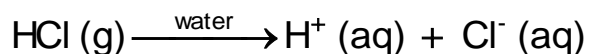


Figure 4.21 These are magnified nylon fibres formed by forcing molten plastic through hundreds of tiny holes.

Properties of covalent compounds

Covalent compounds have the following properties.

- As simple molecular substances, they are usually gases, liquids or solids with low melting and boiling points. The melting points are low because of the weak intermolecular forces of attraction which exist between simple molecules. Giant molecular sub-stances have higher melting points, because the whole structure is held together by strong covalent bonds within the giant molecule.
- Generally, they do not conduct electricity when molten or dissolved in water. This is because they do not contain ions. However, some molecules actually react with water to form ions. For example, hydrogen chloride gas produces aqueous hydrogen ions and chloride ions when it dissolves in water:



- Generally, they do not dissolve in water. However, water is an excellent solvent and can interact with and dissolve some covalent molecules better than others.

Allotropy

When an element can exist in more than one physical form in the same state it is said to exhibit **allotropy** (or polymorphism). Each of the different physical forms is called an **allotrope**. Allotropy is actually quite a common feature of the elements in the periodic table. Some examples of elements which show allotropy are sulphur, tin, iron and carbon.

Allotropes of carbon

Carbon is a non-metallic element which exists in more than one solid structural form. Its allotropes are called **graphite** and **diamond**. Each of the allotropes has a different structure (Figures 4.22 and 4.23) and so the allotropes exhibit different physical properties (Table 4.2). The different physical properties that they exhibit lead to the allotropes being used in

different ways (Table 4.3).

Table 4.2 Physical properties of graphite and diamond.

Property	Graphite	Diamond
Appearance	A dark grey, shiny solid	A colourless transparent crystal which sparkles in light
Electrical conductivity	Conducts electricity	Does not conduct electricity
Hardness	A soft material with a slippery feel	A very hard substance
Density/g·cm ⁻³	2.25	3.51

Table 4.3 Uses of graphite and diamond.

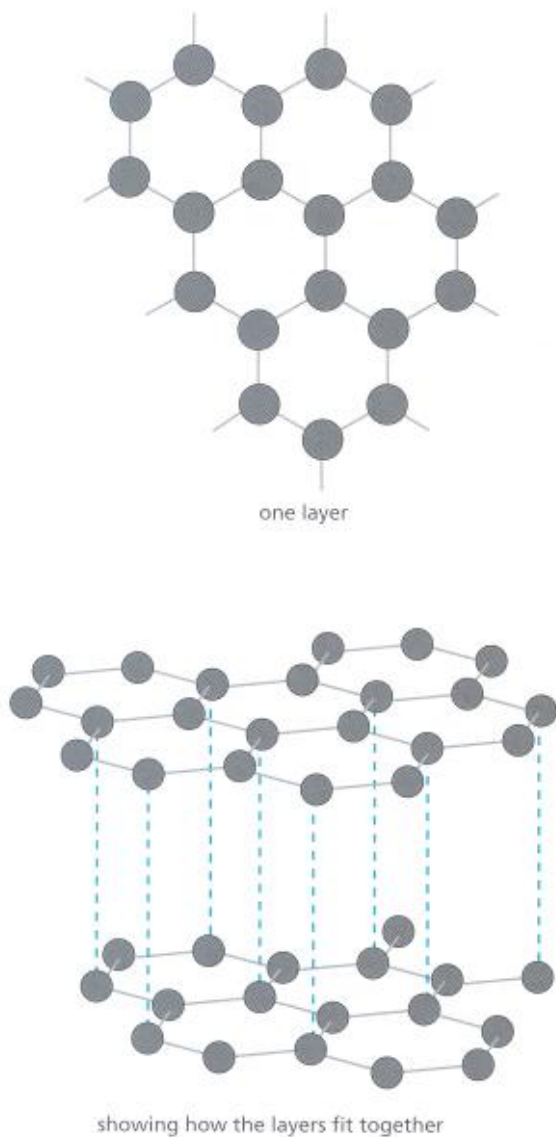
Graphite	Diamond
Pencils	Jewellery
Electrodes	Glass cutters
Lubricant	Diamond-studded saws
	Drill bits
	Polishers

Graphite

Figure 4.22a shows the structure of graphite. This is a layer structure. Within each layer each carbon atom is bonded to three others by strong covalent bonds. Each layer is therefore a giant molecule. Between these layers there are weak forces of attraction (van der Waals' forces) and so the layers will pass over each other easily.

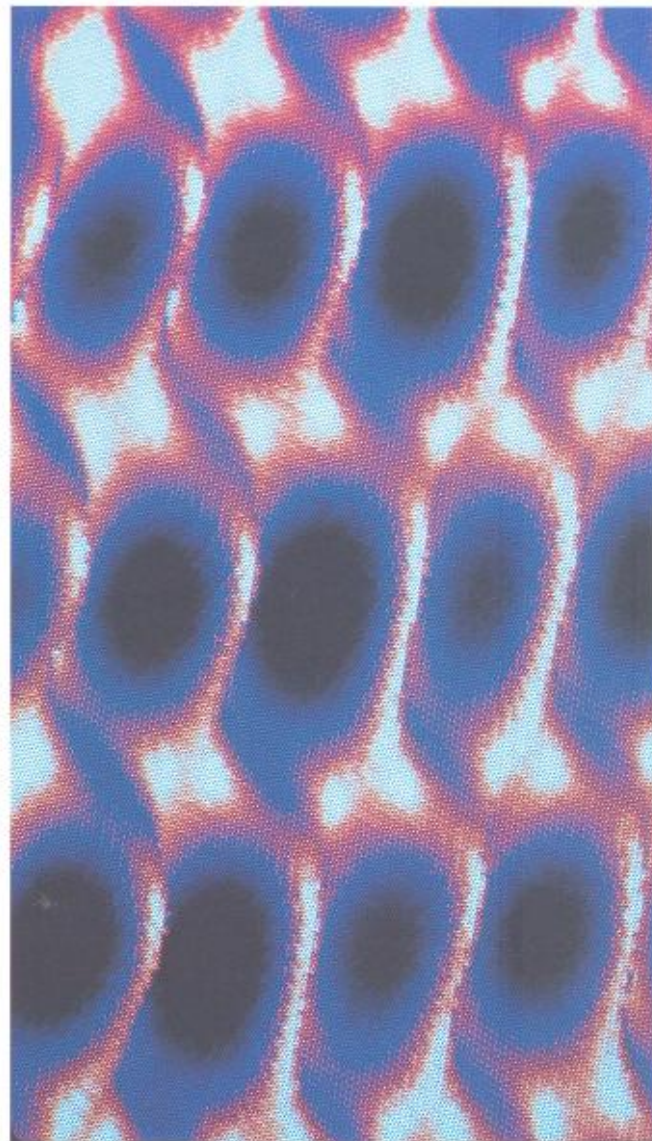
With only three covalent bonds formed between carbon atoms within the layers, an unbonded electron is present on each carbon atom. These 'spare' (or **delocalised**) electrons form electron clouds between the layers and it is because of these spare electrons that graphite conducts electricity.

In recent years a set of interesting compounds known as **graphitic compounds** have been developed. In these compounds different atoms have been fitted in between the layers of carbon atoms to produce a substance with a greater electrical conductivity than pure graphite. Graphite is also used as a component in certain sports equipment, such as tennis and squash rackets.



a A portion of the graphite structure.

Figure 4.22



b A piece of graphite as seen through a scanning tunnelling microscope.

Diamond

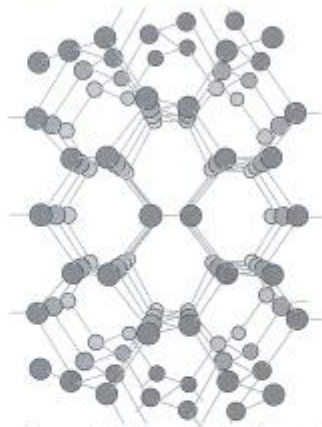
Figure 4.23 shows the diamond structure. Each of the carbon atoms in the giant structure is covalently bonded to four others. They form a tetrahedral arrangement. This bonding scheme gives rise to a very rigid, three-dimensional structure and accounts for the extreme hardness of the substance. All the outer energy level electrons of the carbon atoms are used to form covalent bonds, so there are no electrons available to enable diamond to conduct electricity.

It is possible to manufacture both allotropes of carbon. Diamond is made by heating graphite to about 300°C at very high pressures. Diamond made by this method is known as industrial diamond. Graphite can be made by heating a mixture of coke and sand at a very high temperature in an electric arc furnace for about 24 hours.

The various uses of graphite and diamond result from their differing properties (Figure 4.24).



a small part of the structure



a view of a much larger part of the structure

a The structure of diamond.



b The Regent Diamond has been worn by Queen Elizabeth II.

Figure 4.23

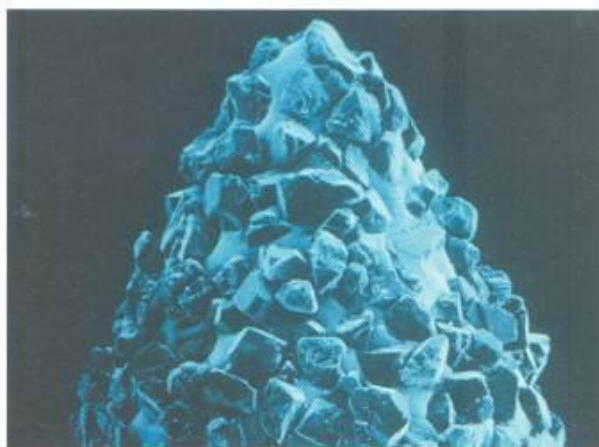


Figure 4.24 Uses of graphite (as a pencil and in a squash racket) and diamond (as a toothed saw to cut marble and on a dentist's drill).

Buckminsterfullerene — an unusual form of carbon

In 1985 a new form of carbon was obtained by Richard Smalley and Robert Curl of Rice University, Texas. It was formed by the action of a laser beam on a sample of graphite. The structure of buckminsterfullerene can be seen in Figure 4.25.

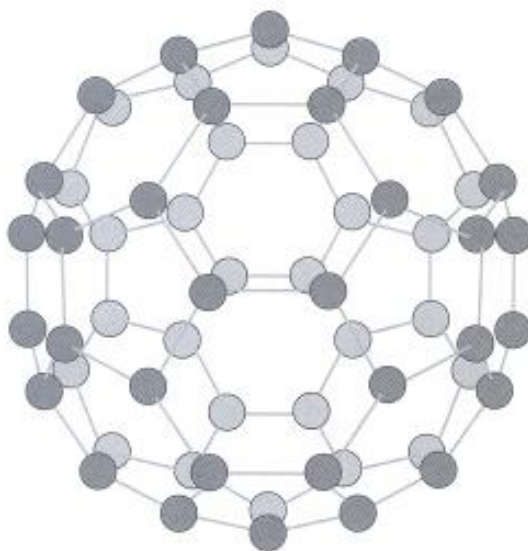


Figure 4.25 Buckminsterfullerene – a 'bucky ball' (C_{60}).

This spherical structure is composed of 60 carbon atoms covalently bonded together. Further spherical forms of carbon, 'bucky balls', containing 70, 72 and 84 carbon atoms have been identified and it has led to a whole new branch of inorganic carbon chemistry. It is thought that this type of molecule exists in chimney soot. Chemists have suggested that due to the large surface area of the bucky balls they may have uses as catalysts (Chapter 11). Also they may have uses as superconductors

Buckminsterfullerene is named after an American architect, Buckminster Fuller, who built complex geometrical structures (Figure 4.26).

Question

1. Explain the difference between ionic and covalent bonding. Discuss in what ways the electron structure of the noble gases is important in both of these theories of bonding.

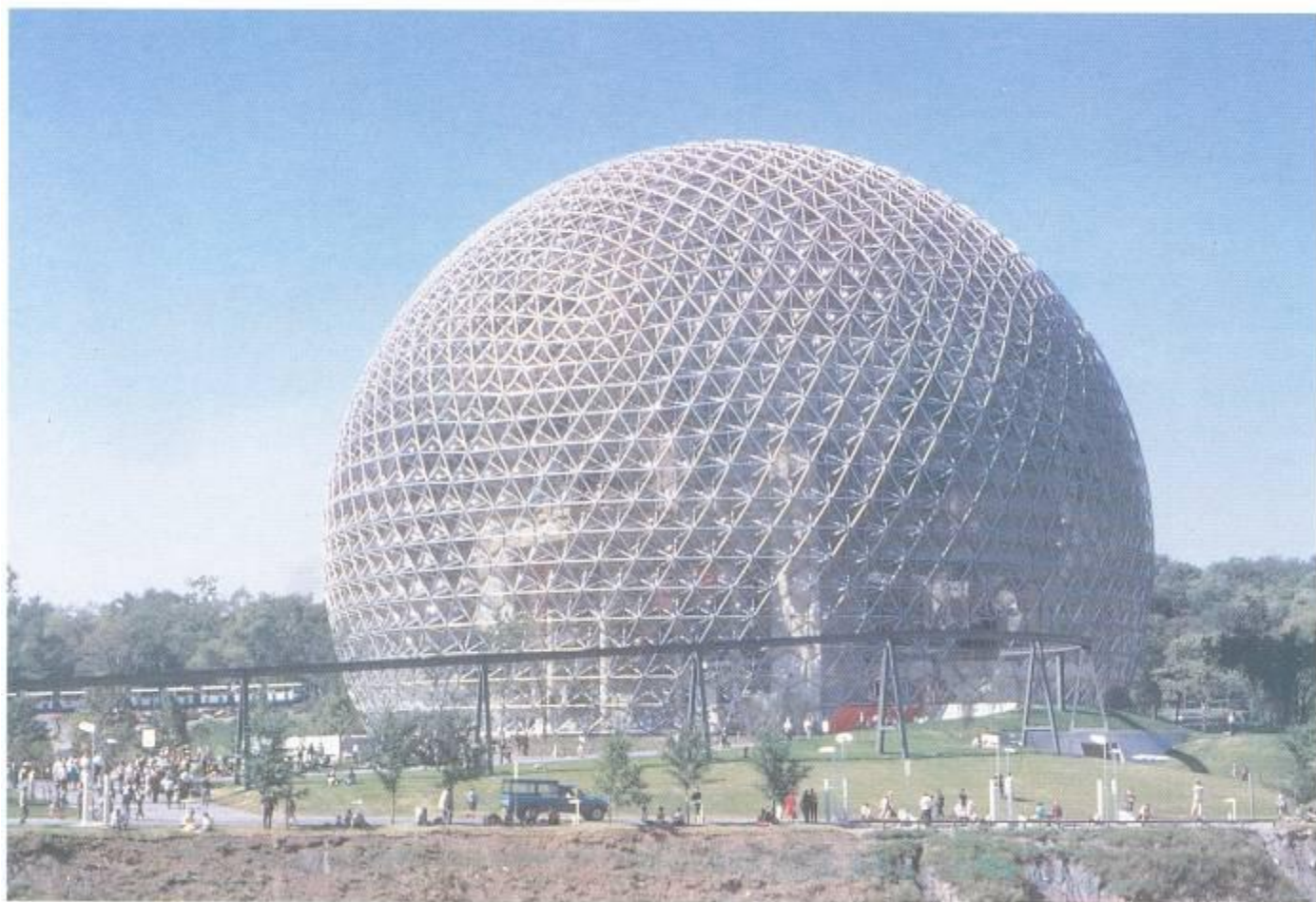


Figure 4.26 C_{60} has a structure similar to a football and to the structure of the dome at the Expo in Montreal, Canada.

Glasses and ceramics

Glasses

Glasses are irregular giant molecular structures held together by strong covalent bonds. Glass can be made by heating silicon(iv) oxide with other substances until a thick viscous liquid is formed. As this liquid cools, the atoms present cannot move freely enough to return to their arrangement within the pure silicon(iv) oxide structure. Instead they are forced to form a disordered arrangement as shown in Figure 4.27. Glass is called a **supercooled liquid**.

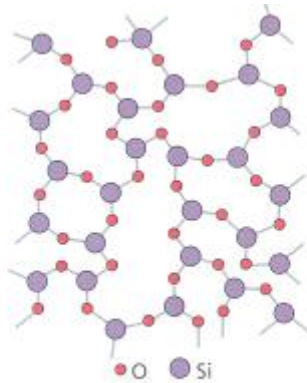


Figure 4.27 Two-dimensional structure of silicon(iv) oxide.

The glass used in bottles and windows is **soda glass**. This type of glass is made by heating a mixture of sand (silicon (IV) oxide), soda (sodium carbonate) and lime (calcium oxide). Pyrex is a borosilicate glass (Figure 4.28). It is made by incorporating some boron oxide into the silicon (IV) oxide structure so that silicon atoms are replaced by boron atoms. This type of glass is tougher than soda glass and more resistant to temperature changes. It is, therefore, used in the manufacture of cooking utensils and laboratory glassware.



Figure 4.28 This glassware is made from Pyrex.

Ceramics

The word **ceramic** comes from the Greek word *keramos* meaning pottery or 'burnt stuff'. Clay dug from the ground contains a mixture of several materials. The main one is a mineral called kaolinite, $\text{Al}_2\text{Si}_2\text{O}_5(\text{OH})_4$, in which the atoms are arranged in layers in a giant structure. While wet, the clay can be moulded because the kaolinite crystals move over one another. However, when it is dry the clay becomes rigid because the crystals stick together.

During firing in a furnace, the clay is heated to a temperature of 1000°C . A complicated series of chemical changes take place, new minerals are formed and some of the substances in the clay react to form a type of glass. The material produced at the end of the firing, the ceramic, consists of many minute mineral crystals bonded together with glass.

Modern ceramic materials now include zirconium oxide (ZrO_2), titanium carbide (TiC), and silicon nitride (SiN). There are now many more uses of these new ceramic materials. For example, vehicle components such as ceramic bearings do not need lubrication — even at high speeds. In space technology, ceramic tiles protect the Space Shuttle from intense heat during its re-entry into the Earth's atmosphere (Figure 4.29).

Questions

1. Draw up a table to summarise the properties of the different types of substances you have met in this chapter. Your table should include examples from ionic substances, covalent substances (simple and giant), ceramics and glasses.
2. Use your research skills, including suitable websites, to discover details of recently developed bioceramics as well as ceramics used as superconductors.



Figure 4.29 The ceramic tiles protect the Space Shuttle from re-entry temperatures of 1500°C . Each tile has fibres coated with silica.

Metallic bonding

Another way in which atoms obtain a more stable electron structure is found in metals. The electrons in the outer energy level of the atom of a metal move freely throughout the structure (they are **delocalised** forming a mobile 'sea' of electrons (Figure 4.30)). When the metal atoms lose these electrons, they become positive ions. Therefore, metals consist of positive ions embedded in moving clouds electrons. The negatively charged electrons attract all the positive metal ions and bond them together with strong electrostatic forces of attraction as a single unit. This is the **metallic bond**.

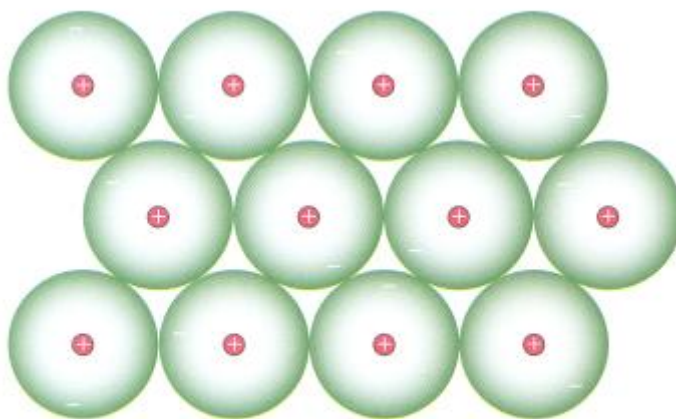


Figure 4.30 Metals consist of positive ions surrounded by a 'sea' of electrons.

Properties of metals

Metals have the following properties.

- They usually have high melting and boiling points due to the strong attraction between the positive metal ions and the mobile 'sea' of electrons.
- They conduct electricity due to the mobile electrons within the metal structure. When a metal is connected in a circuit, the electrons move towards the positive terminal while at the same time electrons are fed into the other end of the metal from the negative terminal.

- They are malleable and ductile. Unlike the fixed bonds in diamond, metallic bonds are not rigid but are still strong. If a force is applied to a metal, rows of ions can slide over one another. They reposition themselves and the strong bonds re-form as shown in Figure 4.31. Malleable means that metals can be hammered into different shapes. Ductile means that the metals can be pulled out into thin wires.
- They have high densities because the atoms are very closely packed in a regular manner as can be seen in Figure 4.32. Different metals show different types of packing and in doing so they produce the arrangements of ions shown in Figure 4.33.

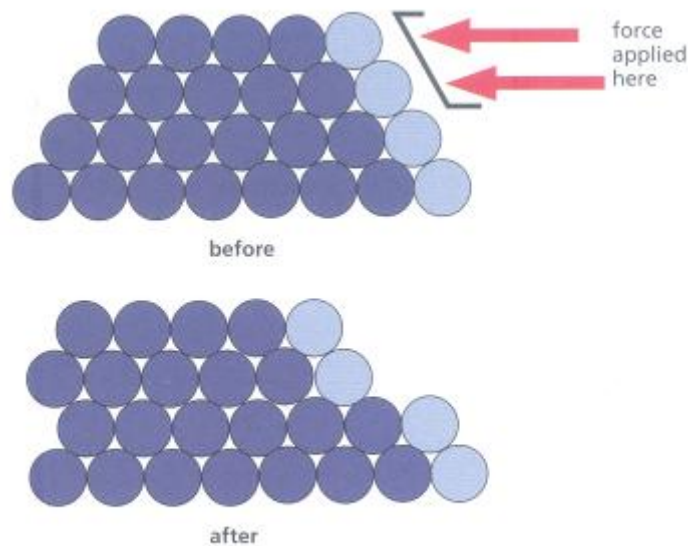


Figure 4.31 The positions of the positive ions in a metal before and after a force has been applied.

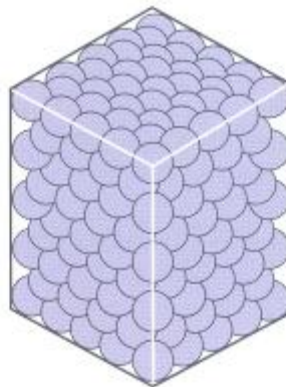


Figure 4.32 Arrangement of ions in the crystal lattice of a metal.

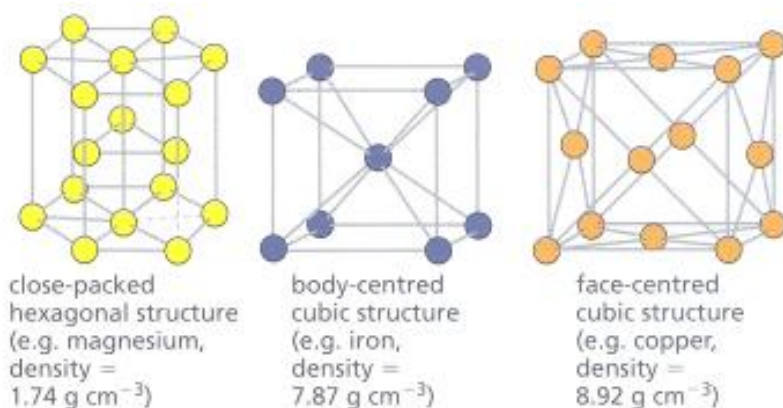


Figure 4.33 Relating different structures to the density of metal.

Questions

1. Explain the terms:
 - a. malleable
 - b. ductile
2. Explain why metals are able to conduct heat and electricity.
3. Explain why the melting point of magnesium ($649 \text{ }^{\circ}\text{C}$) is much higher than the melting point of sodium ($97.9 \text{ }^{\circ}\text{C}$).

Checklist

After studying Chapter 4 you should know and understand the following terms.

Allotropy The existence of an element in two or more different forms in the same physical state.

Ceramics Materials such as pottery made from inorganic chemicals by high-temperature processing. Other modern ceramics include zirconium oxide and silicon nitride.

Covalent bond A chemical bond formed by the sharing of one or more pairs of electrons between two atoms.

Giant ionic structure A lattice held together by the electrostatic forces of attraction between ions.

Giant molecular substance A substance containing thousands of atoms per molecule.

Glass A supercooled liquid which forms a hard, brittle substance that is usually transparent and resistant to chemical attack.

Ionic (electrovalent) bond A strong electrostatic force of attraction between oppositely charged ions.

Intermolecular bonds Attractive forces which act between molecules, for example van der Waals' forces.

Intramolecular bonds Forces which act within a molecule, for example covalent bonds.

Lattice A regular three-dimensional arrangement of atoms/ions in a crystalline solid.

Metallic bond An electrostatic force of attraction between the mobile 'sea' of electrons and the regular array of positive metal ions within the solid metal.

Simple molecular substance These substances possess between one and a few hundred atoms per molecule.

Supercooled liquid One which has cooled below its freezing point without solidification.

Valency The combining power of an atom or group of atoms. The valency of an ion is equal to its charge.

X-ray diffraction A technique often used to study crystal structures.

Bonding and Structure

Additional questions

1. Draw diagrams to show the bonding in each of the following compounds:

a. calcium fluoride (CaF_2)

b. oxygen (O_2)

c. magnesium chloride (MgCl_2)

d. tetrachloromethane (CCl_4).

2. Use the information given in Table 4.1 to work out the formula for:

a. silver oxide

b. zinc chloride

c. potassium sulphate

d. calcium nitrate

e. iron (II) nitrate

f. copper (II) carbonate

g. iron (III) hydroxide

h. aluminium fluoride.

3. Atoms of elements X, Y and Z have 16, 17 and 19 electrons, respectively. Atoms of argon have 18 electrons.

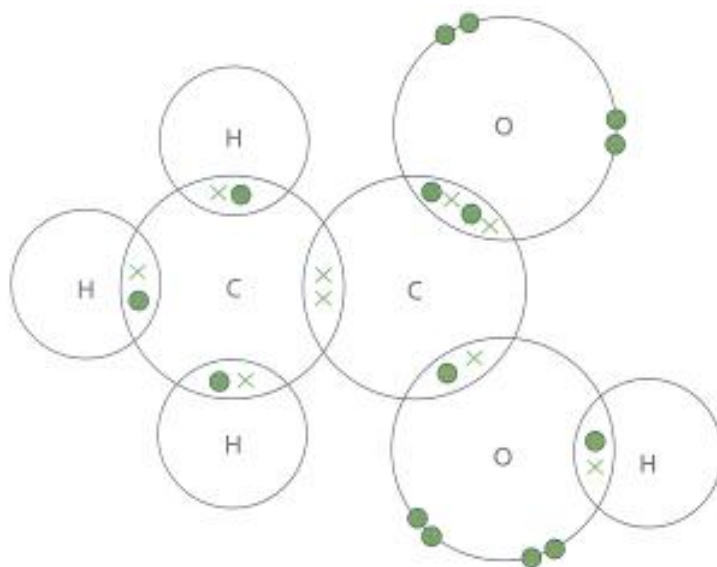
a. Determine the formulae of the compounds formed by the combination of the atoms of the elements:

- (i) X and Z
- (ii) Y and Z
- (iii) X with itself.

b. In each of the cases shown in a(i)—(iii) above, name the type of chemical bond formed.

c. Give two properties you would expect to be shown by the compounds formed in a(ii) and a(iii).

4. The diagram shows the arrangement of the outer electrons only in a molecule of ethanoic acid.



- Name the different elements found in this compound.
- What is the total number of atoms present in this molecule?
- Between which two atoms is there a double covalent bond?
- How many single covalent bonds does each carbon atom have?
- Write a paragraph explaining the sorts of properties you would expect this sort of substance to have.

5. Make a summary table of the properties of substances with covalent structures. Your table should include examples of both simple molecular and giant molecular substances.

6. The elements sodium and chlorine react together to form the compound sodium chloride, which has a giant ionic lattice structure.

a. What type of structure do the elements (i) sodium and (ii) chlorine have?

b. Draw a diagram to represent how the ions are arranged in the crystal lattice of sodium chloride.

c. Explain how the ions are held together in this crystal lattice.

d. Draw diagrams to show how the electrons are arranged in a sodium ion and a chloride ion. (The atomic numbers of sodium and chlorine are 11 and 17, respectively.)

e. Make a table showing the properties of the three substances sodium, chlorine and sodium chloride. Include in your table:

(i) the physical state at room temperature

(ii) solubility in (or reaction with) water

(iii) colour

(iv) electrical conductivity.

7. Explain the following.

- a. Ammonia is a gas at room temperature.
- b. The melting points of sodium chloride and iodine are so different.
- c. Metals generally are good conductors of electricity.
- d. Buckminsterfullerene is an allotrope of carbon.
- e. Metals usually have high melting and boiling points.

8. Discuss the following with reference to diamond and graphite.

a. Diamond is one of the hardest substances known.

b. Graphite is a good lubricating agent.

c. Graphite conducts electricity but diamond does not.